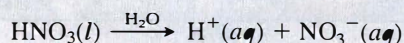


SAMPLE PROBLEM 4.2 Determining the Molarity of H^+ Ions in an Aqueous Solution of an Acid

Problem Nitric acid is a major chemical in the fertilizer and explosives industries. In aqueous solution, each molecule dissociates and the H becomes a solvated H^+ ion. What is the molarity of $\text{H}^+(aq)$ in 1.4 M nitric acid?

Plan We know the molarity of acid (1.4 M), so we just need the formula to find the number of moles of $\text{H}^+(aq)$ present in 1 L of solution.

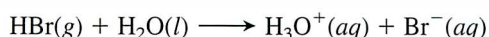
Solution Nitrate ion is NO_3^- , so nitric acid is HNO_3 . Thus, 1 mol of $\text{H}^+(aq)$ is released per mole of acid:



Therefore, 1.4 M HNO_3 contains 1.4 mol of $\text{H}^+(aq)$ per liter and is 1.4 M $\text{H}^+(aq)$.

FOLLOW-UP PROBLEM 4.2 How many moles of $\text{H}^+(aq)$ are present in 451 mL of 3.20 M hydrobromic acid?

Water interacts strongly with many ions, but most strongly with the hydrogen cation, H^+ , a very unusual species. The H atom is a proton surrounded by an electron, so the H^+ ion is just a proton. Because its full positive charge is concentrated in such a tiny volume, H^+ attracts the negative pole of surrounding water molecules so strongly that it actually forms a covalent bond to one of them. We usually show this interaction by writing the aqueous H^+ ion as H_3O^+ (hydronium ion). Thus, to show more accurately what takes place when $\text{HBr}(g)$ dissolves, we should write

**SECTION SUMMARY**

Water plays an active role in dissolving ionic compounds because it consists of polar molecules that are attracted to the ions. When an ionic compound dissolves in water, the ions dissociate from each other and become solvated by water molecules. Because the ions are free to move, their solutions conduct electricity. Water also dissolves many covalent substances with polar bonds. It interacts with some H-containing molecules so strongly it breaks their bonds and dissociates them into $\text{H}^+(aq)$ ions and anions. In water, the H^+ ion is bonded to an H_2O , forming H_3O^+ .

4.2 WRITING EQUATIONS FOR AQUEOUS IONIC REACTIONS

Chemists use three types of equations to represent aqueous ionic reactions: molecular, total ionic, and net ionic equations. As you'll see in the two types of ionic equations, by balancing the atoms, we also balance the charges.

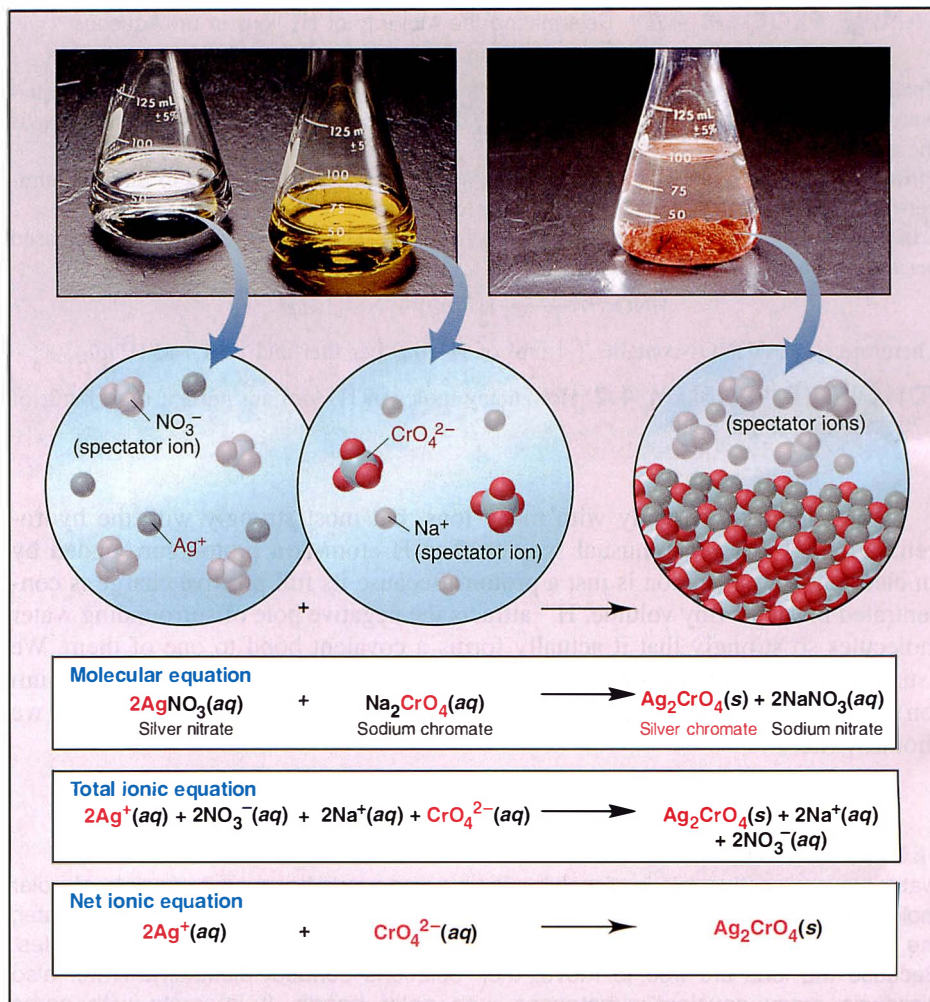
Let's examine a reaction to see what each of these equations shows. When solutions of silver nitrate and sodium chromate are mixed, the brick-red solid silver chromate (Ag_2CrO_4) forms. Figure 4.4 (on the next page) depicts three views of this reaction: the change you would see if you mixed these solutions in the lab, how you might imagine the change at the atomic level among the ions, and how you can symbolize the change with the three types of equations. (The ions that are reacting are shown in red type.)

The **molecular equation** (*top*) reveals the least about the species in solution and is actually somewhat misleading because *it shows all the reactants and products as if they were intact, undissociated compounds*:

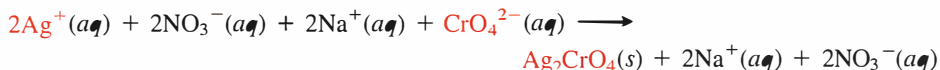


Only by examining the state-of-matter designations (*s*) and (*aq*), can you tell the change that has occurred.

Figure 4.4 An aqueous ionic reaction and its equations. When silver nitrate and sodium chromate solutions are mixed, a reaction occurs that forms solid silver chromate and a solution of sodium nitrate. The photos present the macroscopic view of the reaction, the view the chemist sees in the lab. The blow-up arrows lead to an atomic-scale view, a representation of the chemist's mental picture of the reactants and products. (The pale ions are spectator ions, present for electrical neutrality, but not involved in the reaction.) Three equations represent the reaction in symbols. (The ions that are reacting are shown in red type.) The *molecular equation* shows all substances intact. The *total ionic equation* shows all soluble substances as separate, solvated ions. The *net ionic equation* eliminates the spectator ions to show only the reacting species.



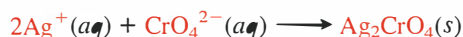
The **total ionic equation** (*middle*) is a much more accurate representation of the reaction because *it shows all the soluble ionic substances dissociated into ions*. Now the $\text{Ag}_2\text{CrO}_4(\text{s})$ stands out as the only undissociated substance:



Notice that charges balance: there are four positive and four negative charges on the left for a net zero charge, and there are two positive and two negative charges on the right for a net zero charge.

Note that $\text{Na}^+(\text{aq})$ and $\text{NO}_3^-(\text{aq})$ appear in the same form on both sides of the equation. They are called **spectator ions** because they are not involved in the actual chemical change. These ions are present as part of the reactants to balance the charge. That is, we can't add an Ag^+ ion without also adding an anion, in this case, NO_3^- ion.

The **net ionic equation** (*bottom*) is the most useful because *it eliminates the spectator ions and shows the actual chemical change taking place*:



The formation of solid silver chromate from silver ions and chromate ions *is* the only change. In fact, if we had originally mixed solutions of potassium chromate, $\text{K}_2\text{CrO}_4(\text{aq})$, and silver acetate, $\text{AgC}_2\text{H}_3\text{O}_2(\text{aq})$, instead of sodium chromate and silver nitrate, the same change would have occurred. Only the

spectator ions would differ— $\text{K}^+(\text{aq})$ and $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$ instead of $\text{Na}^+(\text{aq})$ and $\text{NO}_3^-(\text{aq})$.

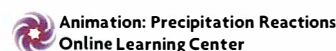
Now, let's apply these types of equations to the three most important types of chemical reactions—precipitation, acid-base, and oxidation-reduction.

SECTION SUMMARY

A molecular equation for an aqueous ionic reaction shows undissociated substances. A total ionic equation shows all soluble ionic compounds as separate, solvated ions. Spectator ions appear unchanged on both sides of the equation. By eliminating them, you see the actual chemical change in a net ionic equation.

4.3 PRECIPITATION REACTIONS

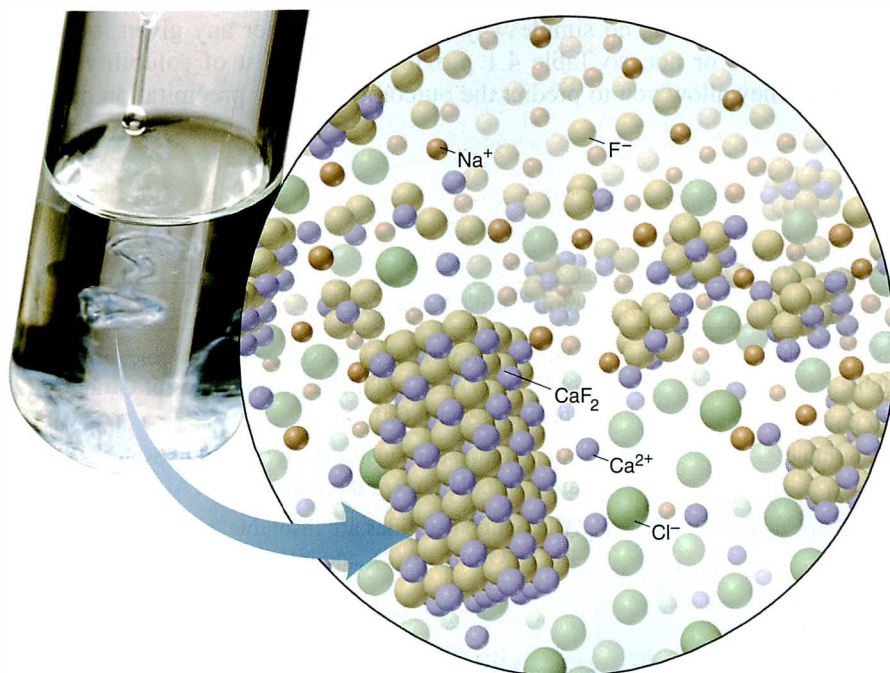
Precipitation reactions are common in both nature and commerce. Many geological formations, including coral reefs, some gems and minerals, and deep-sea structures form, in part, through this type of chemical process. And the chemical industry employs precipitation methods to produce several important inorganic compounds.



The Key Event: Formation of a Solid from Dissolved Ions

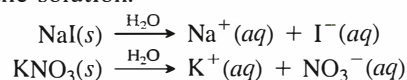
In **precipitation reactions**, two soluble ionic compounds react to form an insoluble product, a **precipitate**. The reaction you just saw between silver nitrate and sodium chromate is an example. Precipitates form for the same reason that some ionic compounds do not dissolve: the electrostatic attraction between the ions outweighs the tendency of the ions to remain solvated and move randomly throughout the solution. When solutions of such ions are mixed, the ions collide and stay together, and the resulting substance “comes out of solution” as a solid, as shown in Figure 4.5 for calcium fluoride. Thus, the key event in a precipitation reaction is *the formation of an insoluble product through the net removal of solvated ions from solution*.

Figure 4.5 The precipitation of calcium fluoride. When an aqueous solution of NaF is added to a solution of CaCl_2 , Ca^{2+} and F^- ions form particles of solid CaF_2 .



Predicting Whether a Precipitate Will Form

If you mix aqueous solutions of two ionic compounds, can you predict if a precipitate will form? Consider this example. When solid sodium iodide and potassium nitrate are each dissolved in water, each solution consists of separated ions dispersed throughout the solution:

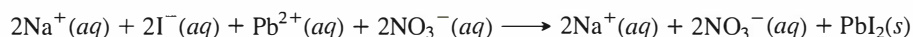


Let's follow three steps to predict whether a precipitate will form:

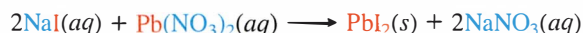
1. *Note the ions present in the reactants.* The reactant ions are

$$\text{Na}^+(aq) + \text{I}^-(aq) + \text{K}^+(aq) + \text{NO}_3^-(aq) \longrightarrow ?$$
2. *Consider the possible cation-anion combinations.* In addition to the two original ones, NaI and KNO₃, which you know are soluble, the other possible cation-anion combinations are NaNO₃ and KI.
3. *Decide whether any of the combinations is insoluble.* A reaction does *not* occur when you mix these starting solutions because all the combinations—NaI, KNO₃, NaNO₃, and KI—are soluble. All the ions remain in solution. (You'll see shortly a set of rules for deciding if a product is soluble or not.)

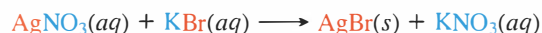
Now, what happens if you substitute a solution of lead(II) nitrate, Pb(NO₃)₂, for the KNO₃? The reactant ions are Na⁺, I[−], Pb²⁺, and NO₃[−]. In addition to the two soluble reactants, NaI and Pb(NO₃)₂, the other two possible cation-anion combinations are NaNO₃ and PbI₂. Lead(II) iodide is insoluble, so a reaction *does* occur because ions are removed from solution (Figure 4.6):



A close look (with color) at the molecular equation shows that *the ions are exchanging partners*:



Such reactions are called double-displacement, or **metathesis** (pronounced *meh-TA-thuh-sis*) **reactions**. Several are important in industry, such as the preparation of silver bromide for the manufacture of black-and-white film:



As we said, there is no simple way to decide whether any given ion combination is soluble or not, so Table 4.1 provides a short list of solubility rules to memorize. They allow you to predict the outcome of many precipitation reactions.

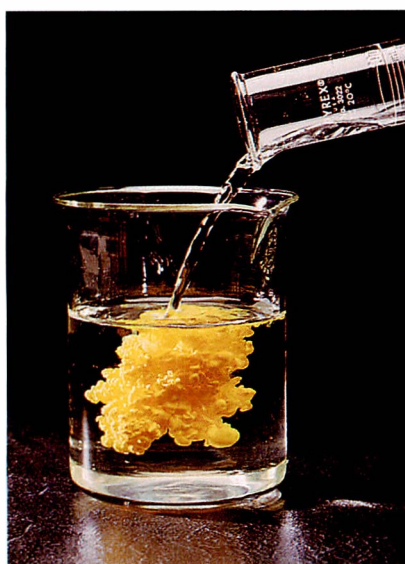


Figure 4.6 The reaction of Pb(NO₃)₂ and NaI. When aqueous solutions of these ionic compounds are mixed, the yellow solid PbI₂ forms.

Table 4.1 Solubility Rules for Ionic Compounds in Water

Soluble Ionic Compounds

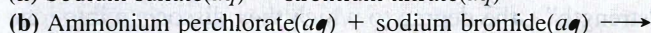
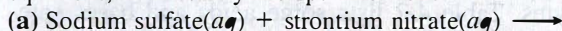
1. All common compounds of Group 1A(1) ions (Li⁺, Na⁺, K⁺, etc.) and ammonium ion (NH₄⁺) are soluble.
2. All common nitrates (NO₃[−]), acetates (CH₃COO[−] or C₂H₃O₂[−]), and most perchlorates (ClO₄[−]) are soluble.
3. All common chlorides (Cl[−]), bromides (Br[−]), and iodides (I[−]) are soluble, *except* those of Ag⁺, Pb²⁺, Cu⁺, and Hg₂²⁺. All common fluorides (F[−]) are soluble, *except* those of Pb²⁺ and Group 2A(2).
4. All common sulfates (SO₄^{2−}) are soluble, *except* those of Ca²⁺, Sr²⁺, Ba²⁺, Ag⁺, and Pb²⁺.

Insoluble Ionic Compounds

1. All common metal hydroxides are insoluble, *except* those of Group 1A(1) and the larger members of Group 2A(2) (beginning with Ca²⁺).
2. All common carbonates (CO₃^{2−}) and phosphates (PO₄^{3−}) are insoluble, *except* those of Group 1A(1) and NH₄⁺.
3. All common sulfides are insoluble *except* those of Group 1A(1), Group 2A(2), and NH₄⁺.

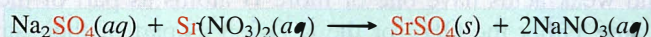
SAMPLE PROBLEM 4.3 Predicting Whether a Precipitation Reaction Occurs; Writing Ionic Equations

Problem Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, total ionic, and net ionic equations, and identify the spectator ions.

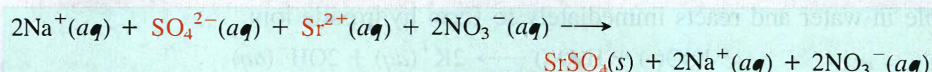


Plan For each pair of solutions, we note the ions present in the reactants, write the cation-anion combinations, and refer to Table 4.1 to see if any are insoluble. For the molecular equation, we predict the products. For the total ionic equation, we write the soluble compounds as separate ions. For the net ionic equation, we eliminate the spectator ions.

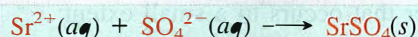
Solution (a) In addition to the reactants, the two other ion combinations are strontium sulfate and sodium nitrate. Table 4.1 shows that strontium sulfate is insoluble, so a reaction *does* occur. Writing the molecular equation:



Writing the total ionic equation:



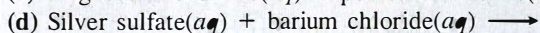
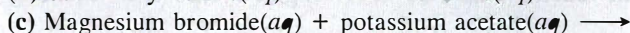
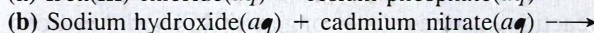
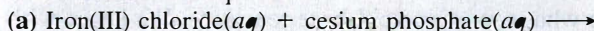
Writing the net ionic equation:



The spectator ions are Na^+ and NO_3^- .

(b) The other ion combinations are ammonium bromide and sodium perchlorate. Table 4.1 shows that all ammonium, sodium, and most perchlorate compounds are soluble, and all bromides are soluble except those of Ag^+ , Pb^{2+} , Cu^+ , and Hg_2^{2+} . Therefore, **no** reaction occurs. The compounds remain dissociated in solution as solvated ions.

FOLLOW-UP PROBLEM 4.3 Predict whether a reaction occurs, and write balanced total and net ionic equations:



SECTION SUMMARY

Precipitation reactions involve the formation of an insoluble ionic compound from two soluble ones. They occur because electrostatic attractions among certain pairs of solvated ions are strong enough to cause their removal from solution. Such reactions can be predicted by noting whether any possible ion combinations are insoluble, based on a set of solubility rules.

4.4 ACID-BASE REACTIONS

Aqueous acid-base reactions involve water not only as solvent but also in the more active roles of reactant and product. These reactions occur in processes as diverse as the biochemical synthesis of proteins, the industrial production of fertilizer, and some of the methods for revitalizing lakes damaged by acid rain.

Obviously, an **acid-base reaction** (also called a **neutralization reaction**) occurs when an acid reacts with a base, but the definitions of these terms and the scope of this reaction class have changed considerably over the years. For our purposes at this point, we'll use definitions that apply to chemicals you commonly encounter in the lab:

- An **acid** is a substance that produces H^+ ions when dissolved in water:

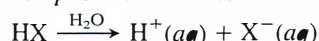


Table 4.2 Strong and Weak Acids and Bases**Acids****Strong**

Hydrochloric acid, HCl
 Hydrobromic acid, HBr
 Hydriodic acid, HI
 Nitric acid, HNO₃
 Sulfuric acid, H₂SO₄
 Perchloric acid, HClO₄

Weak

Hydrofluoric acid, HF
 Phosphoric acid, H₃PO₄
 Acetic acid, CH₃COOH
 (or HC₂H₃O₂)

Bases**Strong**

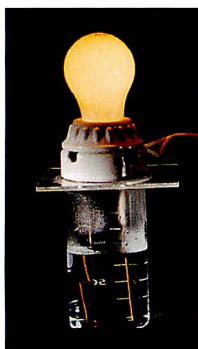
Sodium hydroxide, NaOH
 Potassium hydroxide, KOH
 Calcium hydroxide, Ca(OH)₂
 Strontium hydroxide, Sr(OH)₂
 Barium hydroxide, Ba(OH)₂

Weak

Ammonia, NH₃

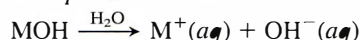


Strong acids and bases are strong electrolytes.



Weak acids and bases are weak electrolytes.

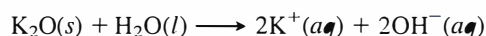
- A **base** is a substance that produces OH[−] ions when dissolved in water.



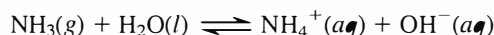
(Other definitions of *acid* and *base* are presented later in this section and again in Chapter 18, along with a fuller meaning of *neutralization*.)

Acids and bases are electrolytes. Table 4.2 lists some acids and bases categorized in terms of their “strength”—the degree to which they dissociate into ions in aqueous solution. In water, *strong acids and strong bases dissociate completely* into ions. Therefore, like soluble ionic compounds, they are *strong* electrolytes and conduct a current well (see left photo in margin). In contrast, *weak acids and weak bases dissociate into ions very little, and most of their molecules remain intact*. As a result, they conduct only a small current and are *weak* electrolytes (see right photo).

Strong and weak acids have one or more H atoms as part of their structure. Strong bases have either the OH[−] or the O^{2−} ion as part of their structure. Soluble ionic oxides, such as K₂O, are strong bases because the oxide ion is not stable in water and reacts immediately to form hydroxide ion:



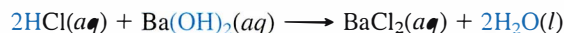
Weak bases, such as ammonia, do not contain OH[−] ions, but they produce them in a reaction with water that occurs to a small extent:



(Note the reaction arrow in the preceding equation. This type of arrow indicates that the reaction proceeds in both directions; we’ll discuss this important idea further in Chapter 17.)

The Key Event: Formation of H₂O from H⁺ and OH[−]

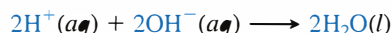
Let’s use ionic equations to see what occurs in acid–base reactions. We begin with the molecular equation for the reaction between the strong acid HCl and the strong base Ba(OH)₂:



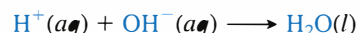
Because HCl and Ba(OH)₂ dissociate completely and H₂O remains undissociated, the total ionic equation is



In the net ionic equation, we eliminate the spectator ions Ba²⁺(aq) and Cl[−](aq) and see the actual reaction:



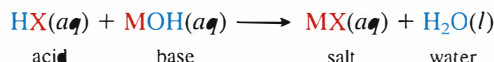
Or



Thus, *the essential change in all aqueous reactions between a strong acid and a strong base is that an H⁺ ion from the acid and an OH[−] ion from the base form a water molecule. In fact, only the spectator ions differ from one strong acid–strong base reaction to another.*

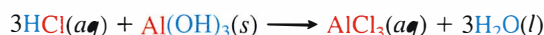
Now it’s easy to understand how these reactions take place: like precipitation reactions, acid–base reactions occur through *the electrostatic attraction of ions and their removal from solution* in the formation of the product. In this case, the ions are H⁺ and OH[−] and the product is H₂O, which consists almost entirely of undissociated molecules. (Actually, water molecules *do* dissociate, but *very* slightly. As you’ll see in Chapter 18, this slight dissociation is very important, but the formation of water in a neutralization reaction nevertheless represents an enormous net removal of H⁺ and OH[−] ions.)

Evaporate the water from the above reaction mixture, and the ionic solid barium chloride remains. An ionic compound that results from the reaction of an acid and a base is called a **salt**. Thus, in a typical aqueous neutralization reaction, *the reactants are an acid and a base, and the products are a salt solution and water*:



The color shows that *the cation of the salt comes from the base and the anion comes from the acid*.

As you can see, acid-base reactions, like precipitation reactions, are metathesis (double-displacement) reactions. The molecular equation for the reaction of aluminum hydroxide, the active ingredient in some antacid tablets, with HCl, the major component of stomach acid, shows this clearly:



Acid-base reactions occur frequently in the synthesis and breakdown of large biological molecules.

SAMPLE PROBLEM 4.4 Writing Ionic Equations for Acid-Base Reactions

Problem Write balanced molecular, total ionic, and net ionic equations for each of the following acid-base reactions and identify the spectator ions:

(a) Strontium hydroxide(aq) + perchloric acid(aq) \longrightarrow

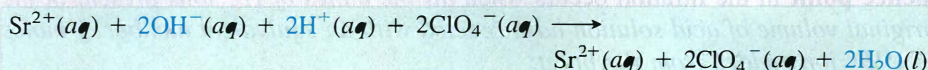
(b) Barium hydroxide(aq) + sulfuric acid(aq) \longrightarrow

Plan All are strong acids and bases (see Table 4.2), so the essential reaction is between H^+ and OH^- . The products are H_2O and a salt solution consisting of the spectator ions. Note that in (b), the salt (BaSO_4) is insoluble (see Table 4.1), so virtually all ions are removed from solution.

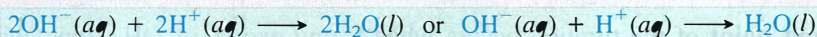
Solution (a) Writing the molecular equation:



Writing the total ionic equation:

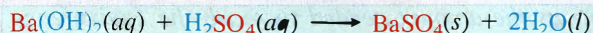


Writing the net ionic equation:

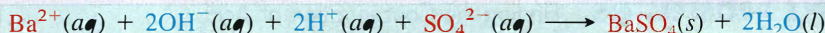


$\text{Sr}^{2+}(aq)$ and $\text{ClO}_4^-(aq)$ are the spectator ions.

(b) Writing the molecular equation:



Writing the total ionic equation:



The net ionic equation is the same as the total ionic equation. This is a precipitation *and* a neutralization reaction. There are **no spectator ions** because all the ions are used to form the two products.

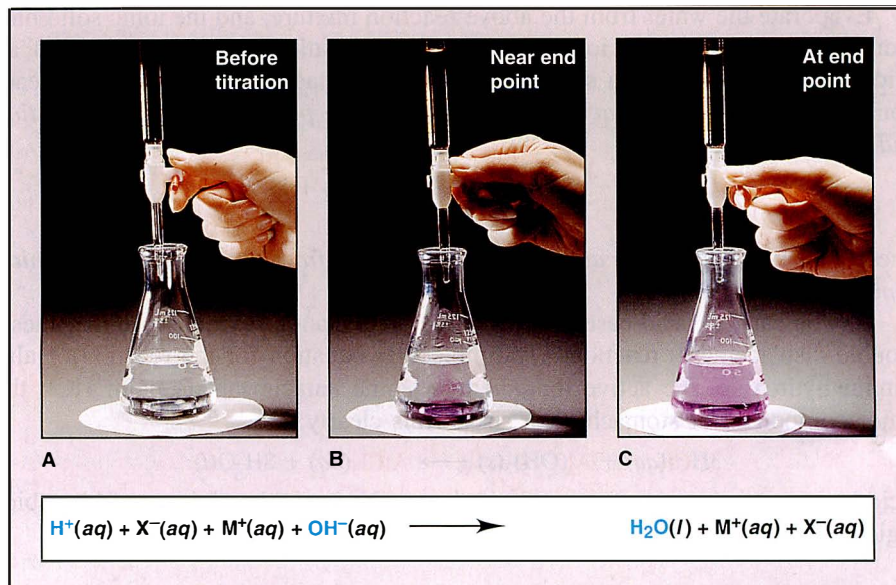
FOLLOW-UP PROBLEM 4.4 Write balanced molecular, total ionic, and net ionic equations for the reaction between aqueous solutions of calcium hydroxide and nitric acid.

Acid-Base Titrations

Chemists study acid-base reactions quantitatively through titrations. In any **titration**, *one solution of known concentration is used to determine the concentration of another solution through a monitored reaction*.

Figure 4.7 An acid-base titration.

A, In this procedure, a measured volume of the unknown acid solution is placed in a flask beneath a buret containing the known (standardized) base solution. A few drops of indicator are added to the flask; the indicator used here is phenolphthalein, which is colorless in acid and pink in base. After an initial buret reading, base (OH^- ions) is added slowly to the acid (H^+ ions). **B**, Near the end of the titration, the indicator momentarily changes to its base color but reverts to its acid color with swirling. **C**, When the end point is reached, a tiny excess of OH^- is present, shown by the permanent change in color of the indicator. The difference between the final buret reading and the initial buret reading gives the volume of base used.



In a typical acid-base titration, a *standardized* solution of base, one whose concentration is *known*, is added slowly to an acid solution of *unknown* concentration (Figure 4.7). A known volume of the acid solution is placed in a flask, and a few drops of indicator solution are added. An *acid-base indicator* is a substance whose color is different in acid than in base. (We examine indicators in Chapters 18 and 19.) The standardized solution of base is added slowly to the flask from a buret. As the titration is close to its end, indicator molecules near a drop of added base change color due to the temporary excess of OH^- ions there. As soon as the solution is swirled, however, the indicator's acidic color returns. The **equivalence point** in the titration occurs when *all the moles of H^+ ions present in the original volume of acid solution have reacted with an equivalent number of moles of OH^- ions added from the buret*:

$$\text{Moles of } \text{H}^+ \text{ (originally in flask)} = \text{moles of } \text{OH}^- \text{ (added from buret)}$$

The **end point** of the titration occurs when a tiny excess of OH^- ions changes the indicator permanently to its color in base. In calculations, we assume this tiny excess is insignificant, and therefore *the amount of base needed to reach the end point is the same as the amount needed to reach the equivalence point*.

SAMPLE PROBLEM 4.5 Finding the Concentration of Acid from an Acid-Base Titration

Problem You perform an acid-base titration to standardize an HCl solution by placing 50.00 mL of HCl in a flask with a few drops of indicator solution. You put 0.1524 M NaOH into the buret, and the initial reading is 0.55 mL. At the end point, the buret reading is 33.87 mL. What is the concentration of the HCl solution?

Plan We must find the molarity of acid from the volume of acid (50.00 mL), the initial (0.55 mL) and final (33.87 mL) volumes of base, and the molarity of base (0.1524 M). First, we balance the equation. We find the volume of base added from the difference in buret readings and use the base's molarity to calculate the amount (mol) of base added. Then, we use the molar ratio from the balanced equation to find the amount (mol) of acid originally present and divide by the acid's original volume to find the molarity.

Solution Writing the balanced equation:



Finding volume (L) of NaOH solution added:

$$\begin{aligned}\text{Volume (L) of solution} &= (33.87 \text{ mL soln} - 0.55 \text{ mL soln}) \times \frac{1 \text{ L}}{1000 \text{ mL}} \\ &= 0.03332 \text{ L soln}\end{aligned}$$

Finding amount (mol) of NaOH added:

$$\begin{aligned}\text{Moles of NaOH} &= 0.03332 \text{ L soln} \times \frac{0.1524 \text{ mol NaOH}}{1 \text{ L soln}} \\ &= 5.078 \times 10^{-3} \text{ mol NaOH}\end{aligned}$$

Finding amount (mol) of HCl originally present: Because the molar ratio is 1:1,

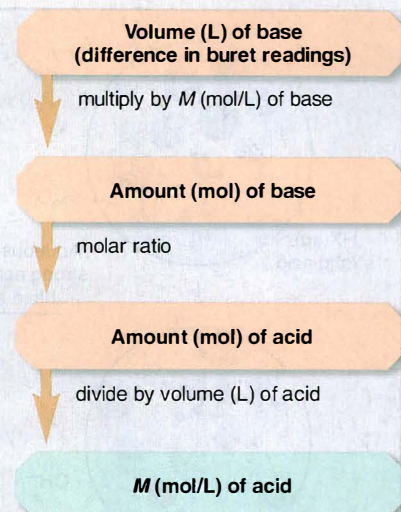
$$\text{Moles of HCl} = 5.078 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 5.078 \times 10^{-3} \text{ mol HCl}$$

Calculating molarity of HCl:

$$\begin{aligned}\text{Molarity of HCl} &= \frac{5.078 \times 10^{-3} \text{ mol HCl}}{50.00 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \\ &= 0.1016 \text{ M HCl}\end{aligned}$$

Check The answer makes sense: a larger volume of less concentrated acid neutralized a smaller volume of more concentrated base. Rounding shows that the moles of H^+ and OH^- are about equal: $50 \text{ mL} \times 0.1 \text{ M H}^+ = 0.005 \text{ mol} = 33 \text{ mL} \times 0.15 \text{ M OH}^-$.

FOLLOW-UP PROBLEM 4.5 What volume of $0.1292 \text{ M Ba(OH)}_2$ would neutralize 50.00 mL of the HCl solution standardized in the preceding sample problem?



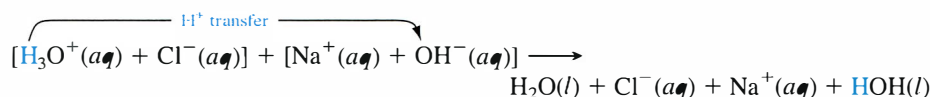
Proton Transfer: A Closer Look at Acid-Base Reactions

We gain deeper insight into acid-base reactions if we look closely at the species in solution. Let's see what takes place when HCl gas dissolves in water. Polar water molecules pull apart each HCl molecule, and the H^+ ion ends up bonded to a water molecule. In essence, HCl *transfers its proton* to H_2O :

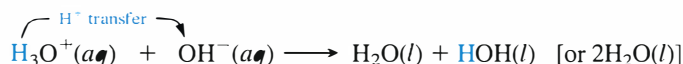


Thus, hydrochloric acid (an aqueous solution of HCl gas) actually consists of solvated H_3O^+ and Cl^- ions.

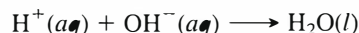
When sodium hydroxide solution is added, the H_3O^+ ion transfers a proton to the OH^- ion of the base (with the product water shown here as HOH):



Without the spectator ions, the transfer of a proton from H_3O^+ to OH^- is obvious:



This net ionic equation is identical with the one we saw earlier (see p. 118),



with the additional H_2O molecule coming from the H_3O^+ . Thus, *an acid-base reaction is a proton-transfer process*. In this case, the Cl^- and Na^+ ions remain in solution, and if the water is evaporated, they crystallize as the salt NaCl. Figure 4.8 (on the next page) shows this process on the atomic level.

In the early 20th century, the chemists Johannes Brønsted and Thomas Lowry realized the proton-transfer nature of acid-base reactions. They defined *an acid as a molecule (or ion) that donates a proton*, and *a base as a molecule (or ion) that accepts a proton*. Therefore, in the aqueous reaction between strong acid and

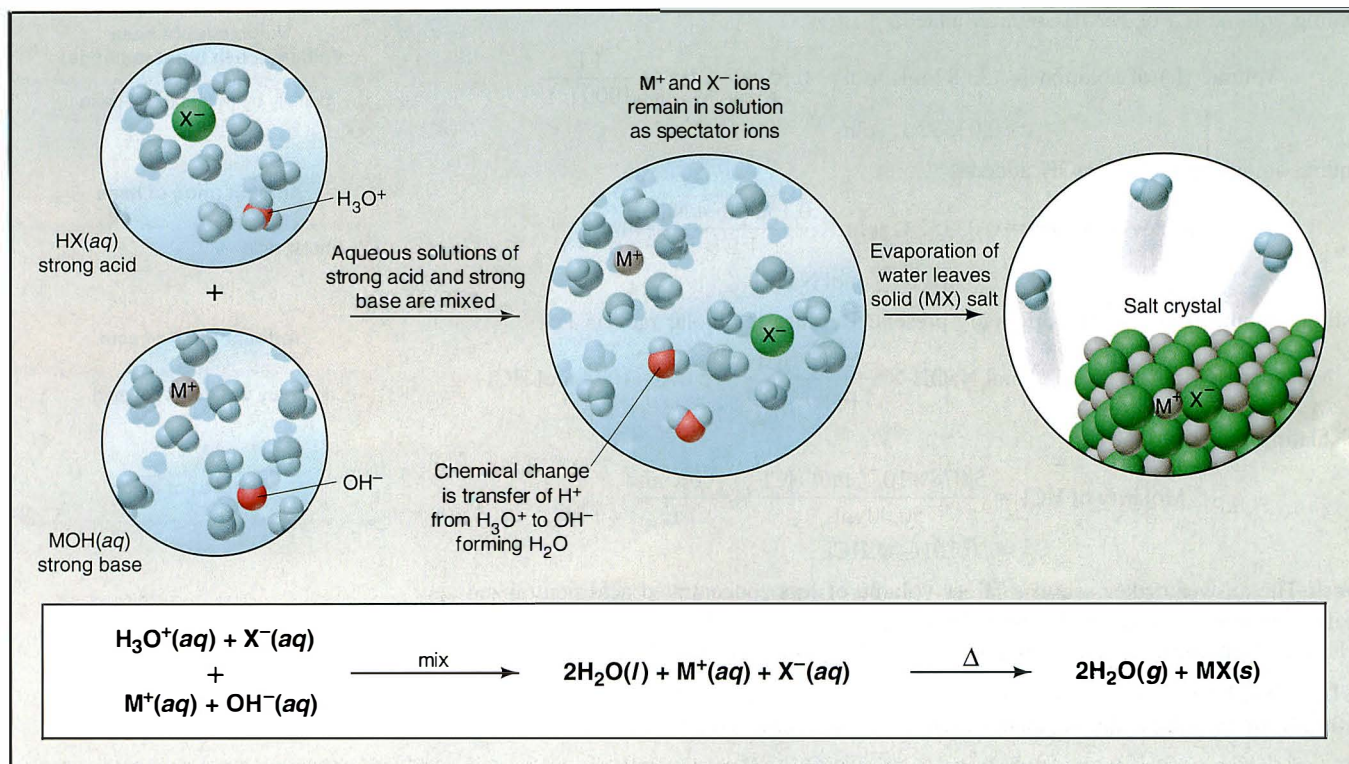
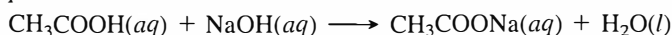


Figure 4.8 An aqueous strong acid–strong base reaction on the atomic scale. When solutions of a strong acid (HX) and a strong base (MOH) are mixed, the H_3O^+ from the acid transfers a proton to the OH^- from the base to form an H_2O molecule. Evaporation of the water leaves the spectator ions, X^- and M^+ , as a solid ionic compound called a *salt*.

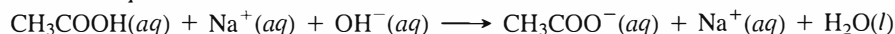
strong base, H_3O^+ ion acts as the acid and OH^- ion acts as the base. Because it ionizes completely, a given amount of strong acid (or strong base) creates an equivalent amount of H_3O^+ (or OH^-) when it dissolves in water. (We discuss the Brønsted-Lowry concept thoroughly in Chapter 18.)

Reactions of Weak Acids Ionic equations are written differently for the reactions of weak acids. When solutions of sodium hydroxide and acetic acid (CH_3COOH) are mixed, the molecular, total ionic, and net ionic equations are

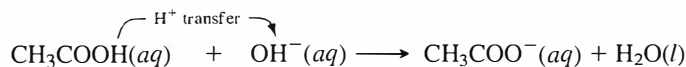
Molecular equation:



Total ionic equation:



Net ionic equation:



Acetic acid dissociates very little because it is a weak acid (see Table 4.2). To show this, *it appears undissociated in both ionic equations*. Note that H_3O^+ does not appear; rather, the proton is transferred from CH_3COOH . Therefore, only $\text{Na}^+(\text{aq})$ is a spectator ion; $\text{CH}_3\text{COO}^-(\text{aq})$ is not.

SECTION SUMMARY

Acid-base (neutralization) reactions occur when an acid (an H^+ -yielding substance) and a base (an OH^- -yielding substance) react and the H^+ and OH^- ions form a water molecule. Strong acids and bases dissociate completely in water; weak acids and bases dissociate slightly. In a titration, a known concentration of one reactant is used to determine the concentration of the other. An acid-base reaction can also be viewed as the transfer of a proton from an acid to a base. Because weak acids dissociate very little, equations involving them show the acid as an intact molecule.

4.5 OXIDATION-REDUCTION (REDOX) REACTIONS

Redox reactions are the third and, perhaps, most important type of chemical process. They include the formation of a compound from its elements (and vice versa), all combustion reactions, the reactions that generate electricity in batteries, the reactions that produce cellular energy, and many others. In this section, we examine the process and introduce some essential terminology.

The Key Event: Movement of Electrons Between Reactants

In **oxidation-reduction** (or **redox**) reactions, the key chemical event is the *net movement of electrons from one reactant to the other*. This movement of electrons occurs *from the reactant (or atom in the reactant) with less attraction for electrons to the reactant (or atom) with more attraction for electrons*.

Such movement of electron charge occurs in the formation of both ionic and covalent compounds. As an example, let's reconsider the flashbulb reaction (see Figure 3.7), in which an ionic compound, MgO, forms from its elements:

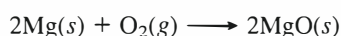
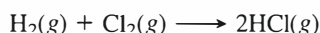


Figure 4.9A shows that during the reaction, each Mg atom loses two electrons and each O atom gains them; that is, two electrons move from each Mg atom to each O atom. This change represents a **transfer of electron charge** away from each Mg atom toward each O atom, resulting in the formation of Mg^{2+} and O^{2-} ions. The ions aggregate and form an ionic solid.

During the formation of a covalent compound from its elements, there is again a net movement of electrons, but it is more of a **shift** in electron charge than a full transfer. Thus, *ions do not form*. Consider the formation of HCl gas:



To see the electron movement here, compare the electron charge distributions in the reactant bonds and in the product bonds. As Figure 4.9B shows, H_2 and Cl_2 molecules are each held together by covalent bonds in which the electrons are shared equally between the atoms (the tan shading is symmetrical). In the HCl molecule, the electrons are shared unequally because the Cl atom attracts them more strongly than the H atom does. Thus, in HCl, the H has less electron charge (*blue shading*) than it had in H_2 , and the Cl has more charge (*red shading*) than it had in Cl_2 . In other words, in the formation of HCl, there has been a **relative shift of electron charge** away from the H atom toward the Cl atom. This electron

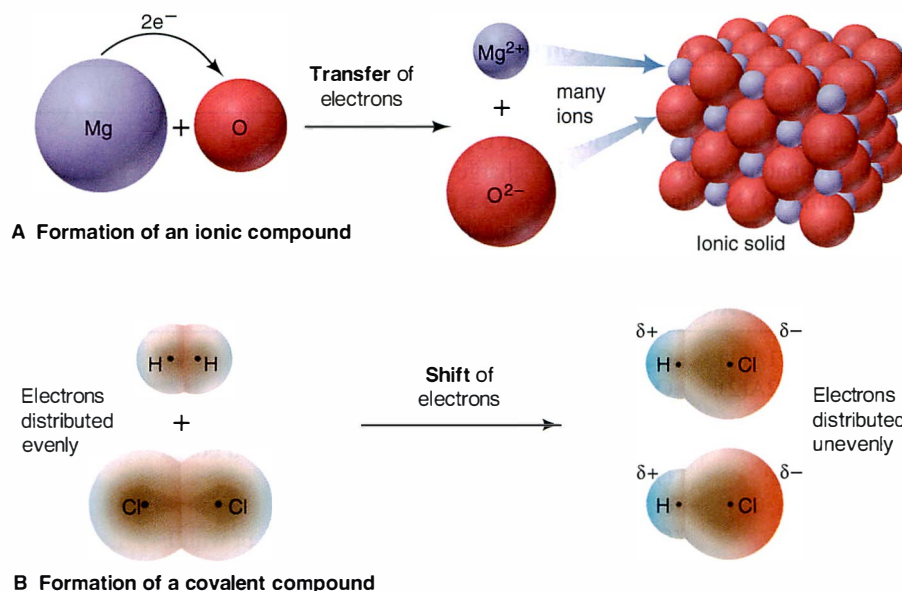


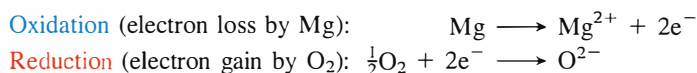
Figure 4.9 The redox process in compound formation. **A**, In forming the ionic compound MgO, each Mg atom transfers two electrons to each O atom. (Note that atoms become smaller when they lose electrons and larger when they gain electrons.) The resulting Mg^{2+} and O^{2-} ions aggregate with many others to form an ionic solid. **B**, In the reactants H_2 and Cl_2 , the electron pairs are shared equally (indicated by even electron density shading). In the covalent product HCl, Cl attracts the shared electrons more strongly than H does. In effect, the H electron shifts toward Cl, as shown by higher electron density (*red*) near the Cl end of the molecule and lower electron density (*blue*) near the H end.

shift is not nearly as extreme as the electron *transfer* during MgO formation. In fact, in some reactions, the net movement of electrons may be very slight, but the reaction is still a redox process.

Some Essential Redox Terminology

Chemists use some important terminology to describe the movement of electrons in oxidation-reduction reactions. **Oxidation** is the *loss* of electrons, and **reduction** is the *gain* of electrons. (The original meaning of *reduction* comes from the process of reducing large amounts of metal ore to smaller amounts of metal, but you'll see shortly why we use the term "reduction" for the act of gaining.)

For example, during the formation of magnesium oxide, Mg undergoes oxidation (electron loss) and O₂ undergoes reduction (electron gain). The loss and gain are simultaneous, but we can imagine them occurring in separate steps:



One reactant acts on the other. Thus, we say that O₂ *oxidizes* Mg, and that O₂ is the **oxidizing agent**, the species doing the oxidizing. Similarly, Mg *reduces* O₂, so Mg is the **reducing agent**, the species doing the reducing.

Note especially that O₂ takes the electrons that Mg loses or, put the other way around, Mg gives up the electrons that O₂ gains. This give-and-take of electrons means that *the oxidizing agent is reduced* because it takes the electrons (and thus gains them), and *the reducing agent is oxidized* because it gives up the electrons (and thus loses them). In the formation of HCl, Cl₂ oxidizes H₂ (H loses some electron charge and Cl gains it), which is the same as saying that H₂ reduces Cl₂. The reducing agent, H₂, is oxidized and the oxidizing agent, Cl₂, is reduced.

Using Oxidation Numbers to Monitor the Movement of Electron Charge

Chemists have devised a useful "bookkeeping" system to monitor which atom loses electron charge and which atom gains it. Each atom in a molecule (or ionic compound) is assigned an **oxidation number (O.N.)**, or *oxidation state*, the charge the atom would have *if* electrons were not shared but were transferred completely. Oxidation numbers are determined by the set of rules in Table 4.3. [Note that an oxidation number has the sign *before* the number (+2), whereas an ionic charge has the sign *after* the number (2+).]

Table 4.3 Rules for Assigning an Oxidation Number (O.N.)

General Rules

1. For an atom in its elemental form (Na, O₂, Cl₂, etc.): O.N. = 0
2. For a monatomic ion: O.N. = ion charge
3. The sum of O.N. values for the atoms in a molecule or formula unit of a compound equals zero. The sum of O.N. values for the atoms in a polyatomic ion equals the ion's charge.

Rules for Specific Atoms or Periodic Table Groups

1. For Group 1A(1): O.N. = +1 in all compounds
2. For Group 2A(2): O.N. = +2 in all compounds
3. For hydrogen: O.N. = +1 in combination with nonmetals
O.N. = -1 in combination with metals and boron
4. For fluorine: O.N. = -1 in all compounds
5. For oxygen: O.N. = -1 in peroxides
O.N. = -2 in all other compounds (except with F)
6. For Group 7A(17): O.N. = -1 in combination with metals, nonmetals (except O),
and other halogens lower in the group

SAMPLE PROBLEM 4.6 Determining the Oxidation Number of an Element

Problem Determine the oxidation number (O.N.) of each element in these compounds:

(a) Zinc chloride (b) Sulfur trioxide (c) Nitric acid

Plan We apply Table 4.3, noting the general rules that the O.N. values in a compound add up to zero, and the O.N. values in a polyatomic ion add up to the ion's charge.

Solution (a) ZnCl_2 . The sum of O.N.s for the monatomic ions in the compound must equal zero. The O.N. of the Zn^{2+} ion is +2. The O.N. of each Cl^- ion is -1, for a total of -2. The sum of O.N.s is $+2 + (-2)$, or 0.

(b) SO_3 . The O.N. of each oxygen is -2, for a total of -6. The O.N.s must add up to zero, so the O.N. of S is +6.

(c) HNO_3 . The O.N. of H is +1, so the O.N.s of the NO_3 group must add up to -1 to give zero for the compound. The O.N. of each O is -2 for a total of -6. Therefore, the O.N. of N is +5.

FOLLOW-UP PROBLEM 4.6 Determine the O.N. of each element in the following:

(a) Scandium oxide (Sc_2O_3) (b) Gallium chloride (GaCl_3)
(c) Hydrogen phosphate ion (d) Iodine trifluoride

The periodic table is a great help in learning the highest and lowest oxidation numbers of most main-group elements, as Figure 4.10 shows:

- For most main-group elements, the A-group number (1A, 2A, and so on) is the *highest* oxidation number (always positive) of any element in the group. The exceptions are O and F (see Table 4.3).
- For main-group nonmetals and some metalloids, the A-group number minus 8 is the *lowest* oxidation number (always negative) of any element in the group.

For example, the highest oxidation number of S (Group 6A) is +6, as in SF_6 , and the lowest is $(6 - 8)$, or -2, as in FeS and other metal sulfides.

Thus, another way to define a redox reaction is one in which *the oxidation numbers of the species change*, and the most important use of oxidation numbers is to monitor these changes:

- If a given atom has a higher (more positive or less negative) oxidation number in the product than it had in the reactant, the reactant species that contains the atom was oxidized (lost electrons). Thus, *oxidation is represented by an increase in oxidation number*.
- If an atom has a lower (more negative or less positive) oxidation number in the product than it had in the reactant, the reactant species that contains the atom was reduced (gained electrons). Thus, *the gain of electrons is represented by a decrease (a "reduction") in oxidation number*.

Figure 4.11 summarizes redox terminology. Oxidation numbers are assigned according to the relative attraction of an atom for electrons, so they are ultimately based on atomic properties, as you'll see in Chapters 8 and 9. (For the remainder of this section, **blue** oxidation numbers represent **oxidation**, and **red** oxidation numbers indicate **reduction**.)

SAMPLE PROBLEM 4.7 Recognizing Oxidizing and Reducing Agents

Problem Identify the oxidizing agent and reducing agent in each of the following:

(a) $2\text{Al}(s) + 3\text{H}_2\text{SO}_4(aq) \longrightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 3\text{H}_2(g)$

(b) $\text{PbO}(s) + \text{CO}(g) \longrightarrow \text{Pb}(s) + \text{CO}_2(g)$

(c) $2\text{H}_2(g) + \text{O}_2(g) \longrightarrow 2\text{H}_2\text{O}(g)$

Plan We first assign an oxidation number (O.N.) to each atom (or ion) based on the rules in Table 4.3. The reactant is the reducing agent if it contains an atom that is oxidized (O.N. increased from left side to right side of the equation). The reactant is the oxidizing agent if it contains an atom that is reduced (O.N. decreased).

		Group number Highest O.N./Lowest O.N.									
		1A	2A	3A	4A	5A	6A	7A			
		+1	+2	+3	+4	+5	+6	+7			
Period	1	H									
	2	Li	Be	B	C	N	O	F			
	3	Na	Mg	Al	Si	P	S	Cl			
	4	K	Ca	Ga	Ge	As	Se	Br			
	5	Rb	Sr	In	Sn	Sb	Te	I			
	6	Cs	Ba	Tl	Pb	Bi	Po	At			
	7	Fr	Ra								

Figure 4.10 Highest and lowest oxidation numbers of reactive main-group elements. The A-group number shows the highest possible oxidation number (O.N.) for a main-group element. (Two important exceptions are O, which never has an O.N. of +6, and F, which never has an O.N. of +7.) For nonmetals (yellow) and metalloids (green), the A-group number minus 8 gives the lowest possible oxidation number.

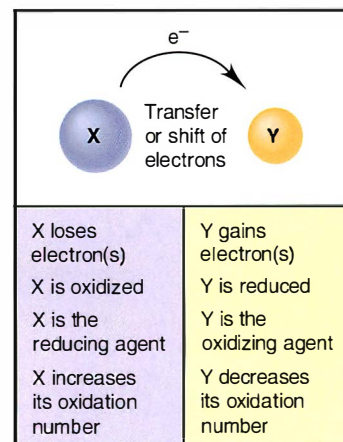
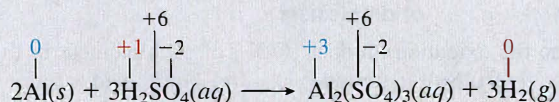


Figure 4.11 A summary of terminology for oxidation-reduction (redox) reactions.

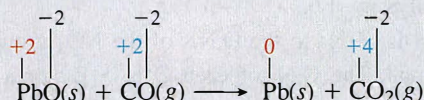
Solution (a) Assigning oxidation numbers:



The O.N. of Al increased from 0 to +3 (Al lost electrons), so Al was oxidized; Al is the reducing agent.

The O.N. of H decreased from +1 to 0 (H gained electrons), so H^+ was reduced; H_2SO_4 is the oxidizing agent.

(b) Assigning oxidation numbers:

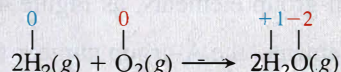


Pb decreased its O.N. from +2 to 0, so PbO was reduced; PbO is the oxidizing agent.

C increased its O.N. from +2 to +4, so CO was oxidized; CO is the reducing agent.

In general, when a substance (such as CO) becomes one with more O atoms (as in CO_2), it is oxidized; and when a substance (such as PbO) becomes one with fewer O atoms (as in Pb), it is reduced.

(c) Assigning oxidation numbers:

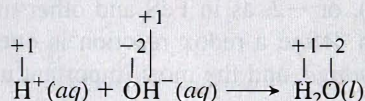


O_2 was reduced (O.N. of O decreased from 0 to -2); O_2 is the oxidizing agent.

H_2 was oxidized (O.N. of H increased from 0 to +1); H_2 is the reducing agent.

Oxygen is always the oxidizing agent in a combustion reaction.

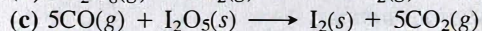
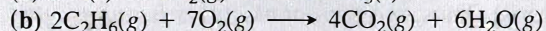
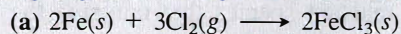
Comment 1. Compare the O.N. values in (c) with those in another common reaction that forms water—the net ionic equation for an acid-base reaction:



Note that the O.N. values remain the same on both sides of the acid-base equation. Therefore, an acid-base reaction is **not** a redox reaction.

2. If a substance occurs in its elemental form on one side of an equation, it can't possibly be in its elemental form on the other side, so the reaction must be a redox process. Notice that elements appear in all three parts of this problem.

FOLLOW-UP PROBLEM 4.7 Identify each oxidizing agent and each reducing agent:



SECTION SUMMARY

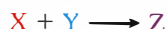
When one reactant has a greater attraction for electrons than another, there is a net movement of electron charge, and a redox reaction takes place. Electron gain (reduction) and electron loss (oxidation) occur simultaneously. The redox process is tracked by assigning oxidation numbers to each atom in a reaction. The species that is oxidized (contains an atom that increases in oxidation number) is the reducing agent; the species that is reduced (contains an atom that decreases in oxidation number) is the oxidizing agent.

4.6 ELEMENTS IN REDOX REACTIONS

As we saw in Sample Problem 4.7, whenever atoms appear in the form of a free element on one side of an equation and as part of a compound on the other, there must have been a change in oxidation state, and the reaction is a redox process.

And, while there are many redox reactions that do *not* involve free elements, we'll focus here on the many others that do. One way to classify these is by comparing the numbers of reactants and products. By that approach, we have three types:

- *Combination reactions*: two or more reactants form one product:



- *Decomposition reactions*: one reactant forms two or more products:

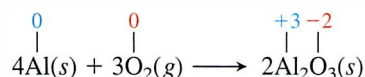


- *Displacement reactions*: the number of substances is the same but atoms (or ions) exchange places:

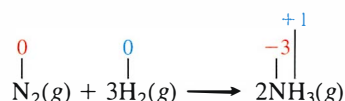


Combining Two Elements Two elements may react to form binary ionic or covalent compounds. Here are some important examples:

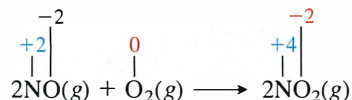
1. *Metal and nonmetal form an ionic compound.* A metal, such as aluminum, reacts with a nonmetal, such as oxygen. The change in O.N.'s shows that the metal is oxidized, so it is the reducing agent; the nonmetal is reduced, so it is the oxidizing agent.



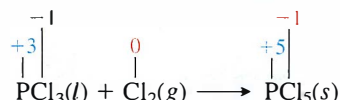
2. *Two nonmetals form a covalent compound.* In one of thousands of examples, ammonia forms from nitrogen and hydrogen in a reaction that occurs in industry on an enormous scale:



Combining Compound and Element Many binary covalent compounds react with nonmetals to form larger compounds. Many nonmetal oxides react with additional O_2 to form “higher” oxides (those with more O atoms in each molecule). For example,

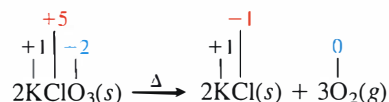


Similarly, many nonmetal halides combine with additional halogen:



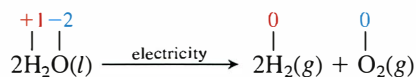
Decomposing Compounds into Elements A decomposition reaction occurs when a reactant absorbs enough energy for one or more of its bonds to break. The energy can take many forms; we'll focus in this discussion on heat and electricity. The products are either elements or elements and smaller compounds. Following are several common examples:

1. *Thermal decomposition.* When the energy absorbed is heat, the reaction is a thermal decomposition. (A Greek *delta*, Δ , above a reaction arrow indicates that heat is required for the reaction.) Many metal oxides, chlorates, and perchlorates release oxygen when strongly heated. Heating potassium chlorate is a method for forming small amounts of oxygen in the laboratory; the same reaction occurs in some explosives and fireworks:

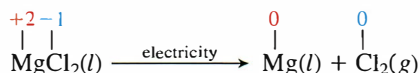


Notice that the lone reactant is the oxidizing *and* the reducing agent.

2. *Electrolytic decomposition.* In the process of *electrolysis*, a compound absorbs electrical energy and decomposes into its elements. Observing the electrolysis of water was crucial in the establishment of atomic masses:



Many active metals, such as sodium, magnesium, and calcium, are produced industrially by electrolysis of their molten halides:



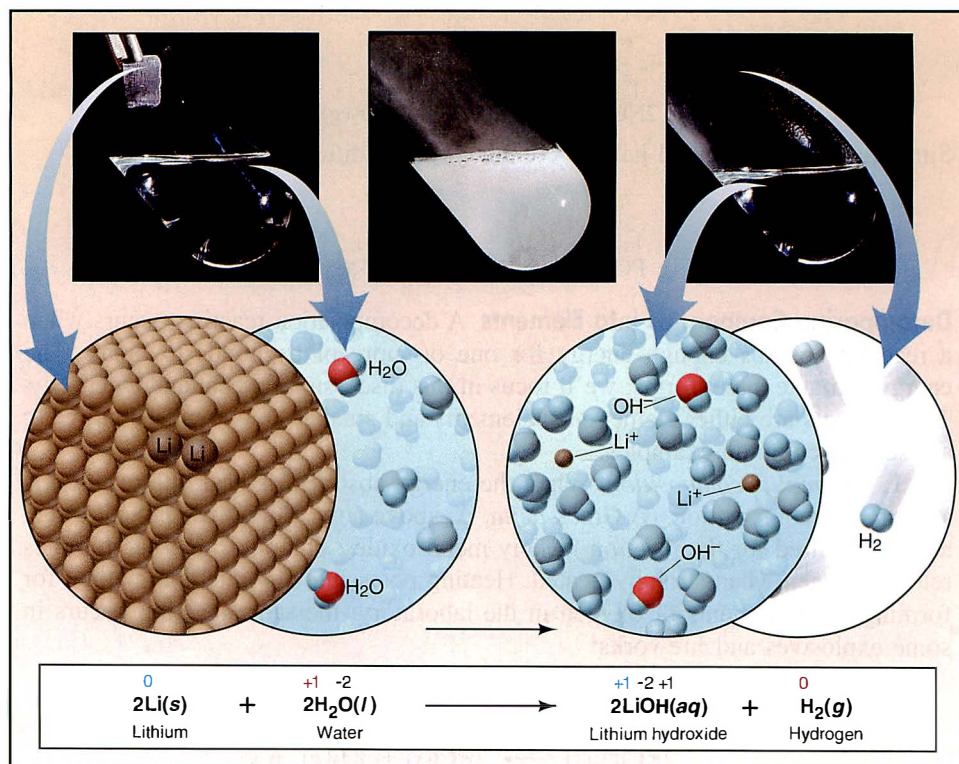
(We'll examine the details of electrolysis in Chapter 21.)

Displacing One Element by Another; Activity Series As we said, displacement reactions have the same number of reactants as products. We mentioned double-displacement (metathesis) reactions in discussing precipitation and acid-base reactions. The other type, *single-displacement* reactions, are all oxidation-reduction processes. They occur when one atom displaces the ion of a different atom from solution. When the reaction involves metals, the atom reduces the ion; when it involves nonmetals (specifically halogens), the atom oxidizes the ion. Chemists rank various elements into activity series—one for metals and one for halogens—in order of their ability to displace one another.

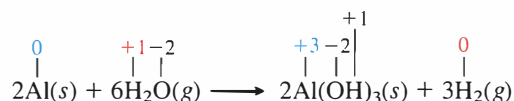
1. *The activity series of the metals.* Metals can be ranked by their ability to displace H_2 (actually reduce H^+) from various sources or by their ability to displace one another from solution.

- *A metal displaces H_2 from water or acid.* The most reactive metals, such as those from Group 1A(1) and Ca, Sr, and Ba from Group 2A(2), displace H_2 from water, and they do so vigorously. Figure 4.12 shows this reaction for

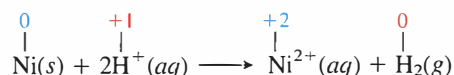
Figure 4.12 An active metal displacing hydrogen from water. Lithium displaces hydrogen from water in a vigorous reaction that yields an aqueous solution of lithium hydroxide and hydrogen gas, as shown on the macroscopic scale (*top*), at the atomic scale (*middle*), and as a balanced equation (*bottom*). (For clarity, the atomic-scale view of water has been greatly simplified, and only water molecules involved in the reaction are colored red and blue.)



lithium. Heat is needed to speed the reaction of slightly less reactive metals, such as Al and Zn, so these metals displace H_2 from steam:



Still less reactive metals, such as nickel and tin, do not react with water but *do* react with acids. Because the concentration of H^+ is higher in acid solutions than in water, H_2 is displaced more easily. Here is the net ionic equation:



Notice that in all such reactions, the metal is the reducing agent (O.N. of metal increases), and water or acid is the oxidizing agent (O.N. of H decreases). The least reactive metals, such as silver and gold, cannot displace H_2 from any source.

- *A metal displaces another metal ion from solution.* Direct comparisons of metal reactivity are clearest in these reactions. For example, zinc metal displaces copper(II) ion from (actually reduces Cu^{2+} in) copper(II) sulfate solution, as the total ionic equation shows:

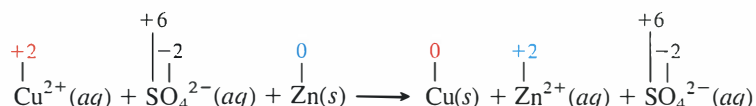
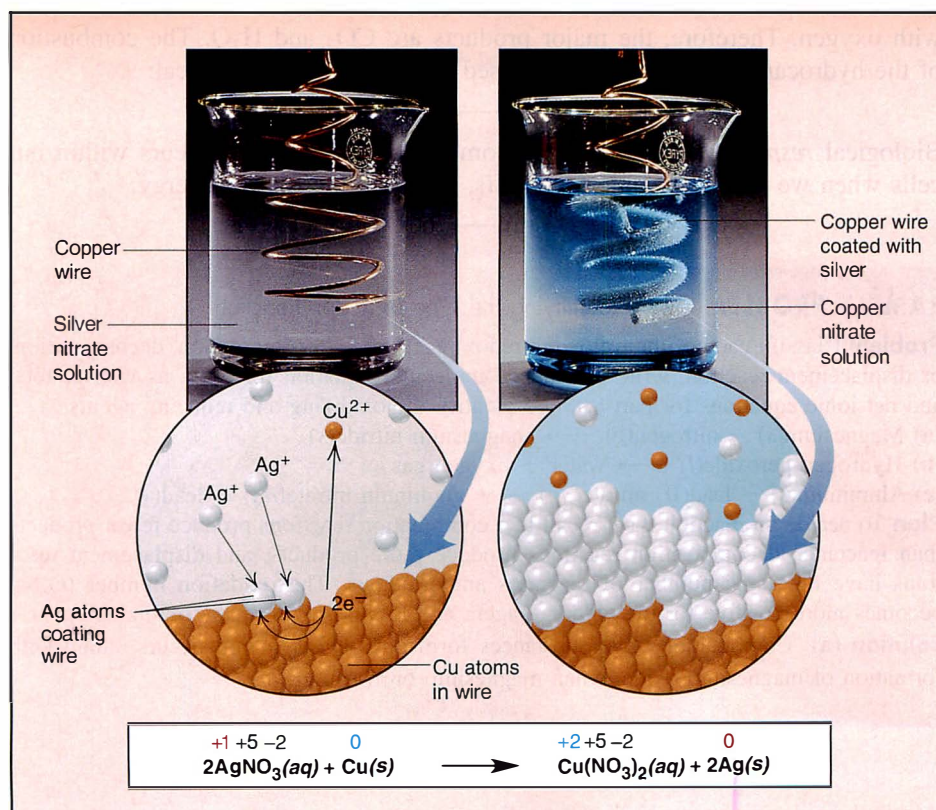


Figure 4.13 demonstrates in atomic detail that copper metal can displace silver ion from solution. Thus, zinc is more reactive than copper, which is more reactive than silver.

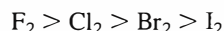


Li	Can displace H ₂ from water
K	
Ba	
Ca	
Na	
Mg	Can displace H ₂ from steam
Al	
Mn	
Zn	
Cr	
Fe	
Cd	Can displace H ₂ from acid
Co	
Ni	
Sn	
Pb	Cannot displace H ₂ from any source
H ₂	
Cu	
Hg	
Ag	
Au	

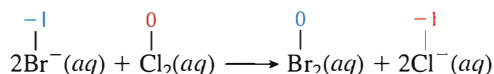
Figure 4.14 The activity series of the metals. This list of metals (and H₂) is arranged with the most active metal (strongest reducing agent) at the top and the least active metal (weakest reducing agent) at the bottom. The four metals below H₂ cannot displace it from any source.

The results of many such reactions between metals and water, aqueous acids, and metal-ion solutions form the basis of the **activity series of the metals**. In Figure 4.14 elements higher on the list are stronger reducing agents than elements lower down; that is, for those that are stable in water, elements higher on the list can reduce aqueous ions of elements lower down. The list also shows whether the metal can displace H₂ (reduce H⁺) and, if so, from which source. Look at the metals in the equations we've just discussed. Note that Li, Al, and Ni lie above H₂, while Ag lies below it; also, Zn lies above Cu, which lies above Ag. The most reactive metals on the list are in Groups 1A(1) and 2A(2) of the periodic table, and the least reactive lie at the right of the transition elements in Groups 1B(11) and 2B(12).

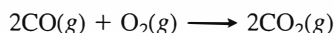
2. *The activity series of the halogens.* Reactivity decreases down Group 7A(17), so we can arrange the halogens into their own activity series:



A halogen higher in the periodic table is a stronger oxidizing agent than one lower down. Thus, chlorine can oxidize bromide ions or iodide ions from solution, and bromine can oxidize iodide ions. Here, chlorine displaces bromine:



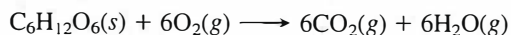
Combustion Reactions *Combustion* is the process of combining with oxygen, often with the release of heat and light, as in a flame. Combustion reactions do not fall neatly into classes based on the number of reactants and products, but *all are redox processes* because elemental oxygen is a reactant:



The combustion reactions that we commonly use to produce energy involve organic mixtures such as coal, gasoline, and natural gas as reactants. These mixtures consist of substances with many carbon-carbon and carbon-hydrogen bonds. During the reaction, these bonds break, and each C and H atom combines with oxygen. Therefore, the major products are CO₂ and H₂O. The combustion of the hydrocarbon butane, which is used in camp stoves, is typical:



Biological *respiration* is a multistep combustion process that occurs within our cells when we “burn” organic foodstuffs, such as glucose, for energy:



SAMPLE PROBLEM 4.8 Identifying the Type of Redox Reaction

Problem Classify each of the following redox reactions as a combination, decomposition, or displacement reaction, write a balanced molecular equation for each, as well as total and net ionic equations for part (c), and identify the oxidizing and reducing agents:

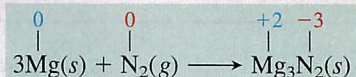
(a) Magnesium(s) + nitrogen(g) \longrightarrow magnesium nitride(s)

(b) Hydrogen peroxide(l) \longrightarrow water + oxygen gas

(c) Aluminum(s) + lead(II) nitrate(aq) \longrightarrow aluminum nitrate(aq) + lead(s)

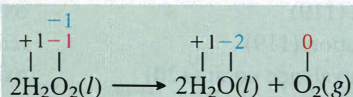
Plan To decide on reaction type, recall that combination reactions produce fewer products than reactants, decomposition reactions produce more products, and displacement reactions have the same number of reactants and products. The oxidation number (O.N.) becomes more positive for the reducing agent and less positive for the oxidizing agent.

Solution (a) **Combination:** two substances form one. This reaction occurs, along with formation of magnesium oxide, when magnesium burns in air:



Mg is the reducing agent; N₂ is the oxidizing agent.

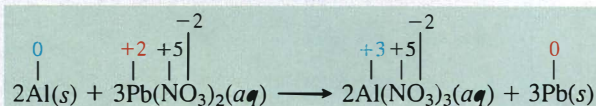
(b) Decomposition: one substance forms two. This reaction occurs within every bottle of this common household antiseptic. Hydrogen peroxide is very unstable and breaks down from heat, light, or just shaking:



H_2O_2 is both the oxidizing *and* the reducing agent. The O.N. of O in peroxides is -1 .

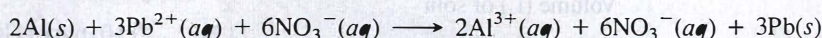
It increases to 0 in O_2 and decreases to -2 in H_2O .

(c) Displacement: two substances form two others. As Figure 4.14 shows, Al is more active than Pb and, thus, displaces it from aqueous solution:

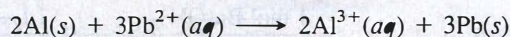


Al is the reducing agent; $\text{Pb}(\text{NO}_3)_2$ is the oxidizing agent.

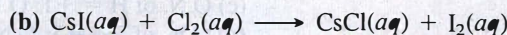
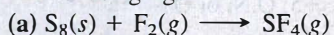
The total ionic equation is



The net ionic equation is



FOLLOW-UP PROBLEM 4.8 Classify each of the following redox reactions as a combination, decomposition, or displacement reaction, write a balanced molecular equation for each, as well as total and net ionic equations for parts (b) and (c), and identify the oxidizing and reducing agents:



SECTION SUMMARY

Any reaction that includes a free element as reactant or product is a redox reaction. In combination reactions, elements combine to form a compound, or a compound and an element combine. Decomposition of compounds by absorption of heat or electricity can form elements or a compound and an element. In displacement reactions, one element displaces another from solution. Activity series rank elements in order of reactivity. The activity series of the metals ranks metals by their ability to displace H_2 from water, steam, or acid, or to displace one another from solution. Combustion typically releases heat and light energy through reaction of a substance with O_2 .

For Review and Reference (Numbers in parentheses refer to pages, unless noted otherwise.)

Learning Objectives

To help you review these learning objectives, the numbers of related sections (§), sample problems (SP), and upcoming end-of-chapter problems (EP) are listed in parentheses.

- Understand how water dissolves an ionic compound compared to a covalent compound and which solution contains an electrolyte; use a compound's formula to find moles of ions (including H^+) in solution (§ 4.1) (SPs 4.1, 4.2) (EPs 4.1–4.19)
- Understand the key events in precipitation and acid-base reactions and use ionic equations to describe them; distinguish between strong and weak acids and bases and calculate an unknown

concentration from a titration (§ 4.3, 4.4) (SPs 4.3–4.5) (EPs 4.20–4.43)

- Understand the key event in the redox process; determine the oxidation number of any element in a compound; identify the oxidizing and reducing agents in a reaction (§ 4.5) (SPs 4.6, 4.7) (EPs 4.44–4.58)

- Identify three important types of redox reactions that include elements: combination, decomposition, displacement (§ 4.6) (SP 4.8) (EPs 4.59–4.73)